#423452
Topic: Oxidation Number

Passage

Assign oxidation number to the underlined elements in each of the following species.

$NaH_2\underset{-}{P}O_4$

Solution

Let X be the oxidation state of P in NaH_2PO_4.

$+1 + 2(1) + X + 4(-2) = +5$

Hence, the oxidation state of P in NaH_2PO_4 is +5.

#423454
Topic: Oxidation Number

Passage

Assign oxidation number to the underlined elements in each of the following species.

$NaH\underset{-}{S}O_4$

Solution

Let X be the oxidation state of S in $NaHSO_4$.

$+1 + 1 + X + 4(2-) = 0$

$X = +6$

Hence, the oxidation state of S in $NaHSO_4$ is +6.

#423455
Topic: Oxidation Number

Passage

Assign oxidation number to the underlined elements in each of the following species.

$H_4\underset{-}{P}_2O_7$

Solution

Let X be the oxidation state of P in $H_4P_2O_7$.

$4(+1) + 2X + 7(-2) = 0$

$X = +5$

Hence, the oxidation state of P in $H_4P_2O_7$ is +5.

#423464
Topic: Oxidation Number

Passage

Assign oxidation number to the underlined elements in each of the following species.

$K_2\underset{-}{Mn}O_4$

Solution

Let x e the oxidation number of Mn in K_2MnO_4.

$2(+1) + X + 4(-2) = 0$

$2 + X - 8 = 0$

$X = 6$

#423468

Topic: Oxidation Number

Assign oxidation number to the underlined elements in each of the following species.

$Ca\underset{-2}{O}$

Solution

Let x be the oxidation number of O in CaO_2.

$2 + 2x = 0$

$x = -1$

#423470

Topic: Oxidation Number

Passage

Assign oxidation number to the underlined elements in each of the following species.

$Na\underset{-}{B}H_4$

Solution

Let x be the oxidation number of B in $NaBH_4$.

$+1 + x + 4(-1) = 0$

$x = +3$

#423474

Topic: Oxidation Number

Passage

Complete the following reactions:

$H_2(g) + M_mO_o(s) \xrightarrow{\Delta}$

Solution

The completed reaction is

$oH_2(g) + M_mO_o(s) \xrightarrow{\Delta} mM(s) + oH_2O(l)$

o moles of H_2 react with 1 mole of $M_mO_o(s)$ to give m moles of $M(s)$ and o moles of $H_2O(l)$

#423477

Topic: Oxidation Number

Passage

Assign oxidation number to the underlined elements in each of the following species.

$H_2\underset{-2}{S}O_7$

Solution

Let X be the oxidation state of S in $H_2S_2O_7$.

$2(+1) + 2X + 7(2-) = 0$

$X = +6$

Hence, the oxidation state of S in $H_2S_2O_7$ is +6.

#423481

Topic: Oxidation Number

Passage

Assign oxidation number to the underlined elements in each of the following species.

$KAl(\underline{S}O_4)_2.12H_2O$

Solution

Let x be the oxidation state of S in $KAl(\underline{S}O_4)_2.\ 12H_2O$.

$+1 + 3 + 2(x + 4(-2)) = 0$

$+4 + 2x - 16 = 0$

$2x = 12$

$x = +6$

#423491

Topic: Oxidation Number

What are the oxidation number of the underlined elements in each of the following and how do you rationalise your results?

$K\underline{I}_3$

Solution

Let X be the oxidation number of I in KI_3

$+1 + 3X = 0$

$X = \frac{1}{3}$

But the oxidation number cannot be fractional.

KI_3 exists as $K^+[I - I \leftarrow I]^-$. A coordinate bond is formed between I_2 molecule and I^- ion. The oxidation number of two I atoms in I_2 molecule is 0 and that of I^- ion is -1.

#423493

Topic: Oxidation Number

Passage

What are the oxidation number of the underlined elements in each of the following and how do you rationalise your results ?

$H_2\underline{S}_4O_6$

Solution

Let X be the oxidation state of S in $H_2S_4O_6$.

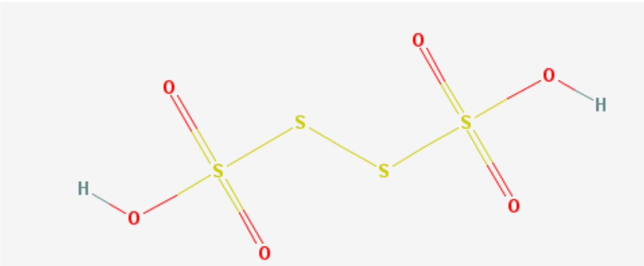
$2(+1) + 4X + 6(2-) = 0$

$X = +\frac{5}{2}$

Hence, the oxidation state of S in $H_2S_4O_6$ is +5/2.

But oxidation number cannot be fractional.

Terminal S atoms have +5 oxidation number and middle S atoms have 0 oxidation number.



#423494
Topic: Oxidation Number

What are the oxidation number of the underlined elements in each of the following and how do you rationalise your results?

$\underline{Fe}_3O_4$

Solution

Let X be the oxidation state of Fe in Fe_3O_4.

$3X + 4(-2) = 0$

$X = +\frac{8}{3}$

Hence, the oxidation state of Fe in Fe_3O_4 is +8/3.

But, oxidation number cannot be fractional.

Fe_3O_4 exists as mixture of FeO and Fe_2O_3, in which Fe has oxidation number of +2 and +3 respectively.

#423498
Topic: Oxidation Number

Passage

What are the oxidation number of the underlined elements in each of the following and how do you rationalise your results ?

$\underline{C}H_3\underline{C}H_2OH$

Solution

For the C atom of methyl group, each H has +1 oxidation number. 3 H atoms of methyl group in total have +3 oxidation number. The oxidation number of C atom of methyl group will be -3 as it will balance the total oxidation number of 3 H atoms.

For the C atom of methylene group, each H has +1 oxidation number and -OH group has -1 oxidation number. The oxidation number of C atom of methylene group will be -1 as it will balance the total oxidation number of 2 H atoms and one -OH group.

Note: The average oxidation number of C atom is as calculated below.

Let X be the oxidation state of C in CH_3CH_2-OH.

$2X + 6(+1) - 2 = 0$

$X = -2$

Hence, the oxidation state of C in CH_3CH_2OH is -2.

#423504
Topic: Oxidation Number

Passage

What are the oxidation number of the underlined elements in each of the following and how do you rationalise your results ?

$\underline{C}H_3\underline{C}OOH$

Solution

(a) By conventional method:

We will determine average oxidation number of C atom.

Let x be the oxidation number of C in CH_3COOH.

The oxidation numbers of H and O are +1 and -2 respectively.

$2X + 4(+1) + 2(-2) = 0$

$x = 0$

Thus, the oxidation number of both the carbon atoms is zero.

(b) According to the structure:

We will determine the oxidation number of each C atom.

(i) For C atom of -COOH group, let X be the oxidation number of this C atom.

This C is attached to one O atom by double bond, one -OH group and on methyl group. The oxidation number of O atom attached by double bond is -2. The oxidation number of -OH group is -1. The C atom of methyl group will not affect the oxidation number of -COOH group.

$x + 1 + 1(-2) + 1(-1) = 0$

$x = +2$

(ii) For methyl carbon atom, let X be the oxidation number of this C atom.

This C atom is attached to 3 H atoms and one -COOH group.

The oxidation number of H atom is +1 and the -COOH group does not affect the oxidation number of C atom of methyl group.

$3(+1) + x + 1(-1) = 0$

$x = -2$

#423513
Topic: Types of redox reactions

Passage

Justify that the following reactions are redox reactions:

$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$

Solution

The oxidation number of Cu decreases from +2 to 0 and that of H increases from 0 to +1. Hence, CuO is reduced to Cu and H_2 is oxidized to H_2O.

Hence, it is a redox reaction.

#423538
Topic: Types of redox reactions

$2K(s) + F_2(g) \rightarrow 2K^+F^-(s)$ is a type of _______ reaction.

A disproportionation

B combustion

C corrosion

$\boxed{\text{D}}$ redox

Solution

The oxidation number of K increases from 0 to +1 and the oxidation number of F2 decreases from 0 to -1. Hence, K is oxidized and F_2 is reduced. Hence, it is redox reaction.

#423548

Topic: Oxidation and reduction - electron transfer concept

Fluorine reacts with ice and results as follows:

$H_2O(s) + F_2(g) \rightarrow HF(g) + HOF(s)$

Justify that this reaction is a redox reaction.

Solution

The oxidation number of F_2 changes from 0 to -1.

Thus, it is reduced. The oxidation number of oxygen changes from -2 to 0. Hence, it is oxidized. Thus, it is a redox reaction.

#423550

Topic: Types of redox reactions

Passage

Complete the following chemical reactions. Classify the below into (a) hydrolysis, (b) redox and (c) hydration reactions.

$MnO_4^-(aq) + H_2O_2(aq) \rightarrow$

Solution

The complete chemical equation is given below:

$MnO_4^-(aq) + 5H_2O_2(aq) + 6H^+ \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5O_2(g)$

This is an example of redox reaction. Mn is reduced and H2O2 is oxidized.

#423554

Topic: Oxidation Number

Calculate the oxidation number of sulphur, chromium and nitrogen in H_2SO_5, $Cr_2O_7^{2-}$ and NO_3^-. Suggest structure of these compounds. Count for the fallacy.

Solution

(a) H_2SO_5 by conventional method.

Let x be the oxidation number of S

$2(+1) + x + 5(-2) = 0$

$x = +8$

+8 Oxidation state of S is not possible as S cannot have oxidation number more than 6. The fallacy is overcomed if we calculate the oxidation number from its structure

$HO - S(O_2) - O - O - H.$

$-1 + X + 2(-2) + 2(-1) + 1 = 0$

$x = +6$

(b) Dichromate ion

Let x be the oxidation number of Cr in dichromate ion

$2x + 7(-2) = -2$

$x = +6$

Hence the oxidation number of Cr in dichromate ion is +6. This is correct and there is no fallacy.

(c) Nitrate ion, by conventional method

Let x be the oxidation number of N in nitrate ion.

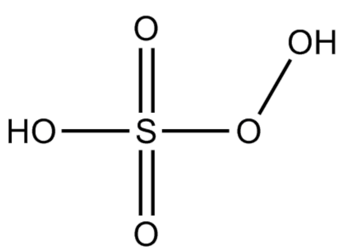
$x + 3(-2) = -1$

From the structure $O^- - N^+(O) - O^-$

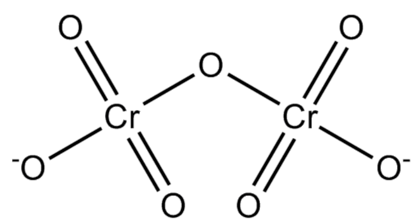
$x + 1(-1) + 1(-2) + 1(-2) = 0$

$x = +5$

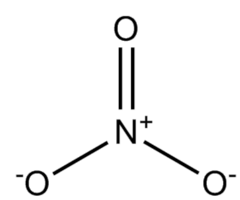
Thus there is no fallacy.



Sulphuric acid



Dichromate



Nitrate

#423557
Topic: Oxidation Number

Passage

Write formulas for the following compounds.

Mercury (II) chloride

Solution

Formula for mercury (II) chloride is $HgCl_2$.

#423560
Topic: Oxidation Number

Passage

Write formulas for the following compounds.

Nickel (II) sulphate

Solution

Formula for Nickel (II) sulphate is $NiSO_4$.

#423562

Topic: Oxidation Number

Passage

Write formulas for the following compounds.

Tin (IV) oxide

Solution

As Tin has $+4$ oxidation state(IV) , and oxidation state of O is -2.

So, there must be 2 O atom so that neutral compound can be formed with correct formula.

So, formula must be SnO_2.

#423565

Topic: Oxidation Number

Passage

Write formulas for the following compounds.

Thallium (I) sulphate

Solution

Formula for thallium(I)sulphate is Tl_2SO_4.

#423566

Topic: Oxidation Number

Passage

Write formulas for the following compounds.

Iron (III) sulphate

Solution

Formula for Iron (III) sulphate is $Fe_2(SO_4)_3$.

#423569

Topic: Oxidation Number

Passage

Write formulas for the following compounds.

Chromium (III) oxide

Solution

Formula for chromium (III) oxide is Cr_2O_3.

#423571

Topic: Oxidation Number

Suggest a list of the substances where carbon can exhibit oxidation states from -4 to +4 and nitrogen from -3 to +5.

Solution

The substances alongwith oxidation states of C are shown below.

Substance	Oxidation number of C
CH_2Cl_2	0
$FC \equiv CF$	+1
$HC \equiv CH$	-1
$CHCl_3$, CO	+2
CH_3Cl	-2
Cl_3C-CCl_3	+3
H_3C-CH_3	-3
CCl_4, CO_2	+4
CH_4	-4

The substances alongwith oxidation states of N are shown below.

Substance	Oxidation number of N
N_2	0
N_2O	+1
N_2H_2	-1
NO	+2
N_2H_4	-2
N_2O_3	+3
NH_3	-3
NO_2	+4
N_2O_5	+5

#423576

Topic: Oxidation and reduction - electron transfer concept

While sulphur dioxide and hydrogen peroxide can act as oxidising as well as reducing agents in their reactions, ozone and nitric acid act only as oxidants. Why ?

Solution

The S atom in SO_2 has +4 oxidation number. The minimum and maximum oxidation numbers of S are -2 and +6 respectively. Hence, in SO_2, S can increase and decrease its oxidation number. Hence, SO_2 is an oxidizing agent as well as reducing agent.

The O atom in hydrogen peroxide has oxidation number of -1. The minimum and maximum oxidation numbers of O are -2 and 0 respectively. Hence, hydrogen peroxide is oxidant as well as reluctant.

In ozone, O atom has oxidation number of 0. It can decrease its oxidation number to -1 or -2 but cannot increase it. Hence ozone is an oxidizing agent.

In nitric acid, N has oxidation number of +5 which is maximum. N can only decrease its oxidation number. Hence, nitric acid is an oxidizing agent.

#423608

Topic: Oxidation and reduction - electron transfer concept

The compound AgF_2 is unstable compound. However, if formed, the compound acts as a very strong oxidising agent. Why ?

Solution

Ag in AgF_2 has +2 oxidation state which is an unstable oxidation state of Ag.

When AgF_2 is formed, silver accepts an electron to form Ag^+.

Hence, silver is reduced and AgF_2 acts as a very strong oxidizing agent.

#423615

Topic: Oxidation Number

Whenever a reaction between an oxidising agent and a reducing agent is carried out, a compound of lower oxidation state is formed if the reducing agent is in excess and a compound of higher oxidation state is formed if the oxidising agent is in excess. Justify this statement giving three illustrations.

Solution

Whenever a reaction between an oxidizing agent and a reducing agent is carried out, a compound of lower oxidation state is formed if the reducing ent is in excess and a compound of higher oxidation state is formed if the oxidizing agent is in excess. Following illustrations justify this.

(i) Oxidizing agent is F_2 and reducing agent is P_4. When excess P_4 reacts with F_2, PF_3 is produced in which P has +3 oxidation number.

$P_4(\text{ excess }) + F_2 \to PF_3$

But if fluorine is in excess, PF_5 is formed in which P has oxidation number of +5.

$P_4 + F_2(\text{ excess }) \to PF_5$

(ii) Oxidizing agent is oxygen and reducing agent is K. When excess K reacts with oxygen, K_2O is formed in which oxygen has oxidation number of -2.

$4K(\text{ excess }) + O_2 \to 2K_2O$

But if oxygen is in excess, then K_2O_2 is formed in which O has oxidation number of -1.

$2K + O_2(\text{ excess }) \to K_2O_2$

(iii) The oxidizing agent is oxygen and the reducing agent is C. When an excess of C reacts with oxygen, CO is formed in which C has +2 oxidation number.

$C(\text{ excess }) + O_2 \to CO$

When excess of oxygen is used, CO_2 is formed in which C has +4 oxidation number.

$C + O_2(\text{ excess }) \to CO_2$

#423663

Topic: Oxidation and reduction - electron transfer concept

$2S_2O_3^{2-}(aq) + I_2(s) \to S_4O_6^{2-}(aq) + 2I^-(aq)$

$S_2O_3^{2-}(aq) + 2Br_2(l) + 5H_2O(l) \to 2SO_4^{2-}(aq) + 4Br^-(aq) + 10H^+(aq)$

Why does the same reductant, thiosulphate react differently with iodine and bromine?

Solution

Bromine is stronger oxidizing agent. Hence, it oxidizes $S_2O_3^{2-}$ to SO_4^{2-}.

Iodine is a weaker oxidizing agent. Hence, it oxidizes $S_2O_3^{2-}$ to $S_4O_6^{2-}$.

#423664

Topic: Oxidation and reduction - electron transfer concept

Justify giving reactions that among halogens, fluorine is the best oxidant and among hydrohalic compounds, hydroiodic acid is the best reductant.

Solution

Fluorine oxidizes chloride ion to chlorine, bromide ion to bromine and iodide ion to iodine respectively.

$F_2 + 2Cl^- \rightarrow 2F^- + Cl_2$

$F_2 + 2Br^- \rightarrow 2F^- + Br_2$

$F_2 + 2I^- \rightarrow 2F^- + I_2$.

Chlorine oxidizes bromide ion to bromine and iodide ion to iodine.

$Cl_2 + Br^- \rightarrow 2Cl^- + Br_2$

$Cl_2 + I^- \rightarrow 2Cl^- + I_2$

Bromine oxidizes iodide ion to iodine.

$Br_2 + I^- \rightarrow 2Br^- + I_2$

But bromine and chlorine cannot oxidize fluoride to fluorine. Hence, fluorine is the best oxidizing agent amongst the halogens. The decreasing order of the oxidizing power of halogens is $F_2 > Cl_2 > Br_2 > I_2$.

HI and HBr can reduce sulphuric acid to sulphur dioxide but HCl and HF cannot. Thus, HI and HBr are stronger reducing agents than HCl and HF.

$2HI + H_2SO_4 \rightarrow I_2 + SO_2 + 2H_2O$

$2HBr + H_2SO_4 \rightarrow Br_2 + SO_2 + 2H_2O$

Iodide ion can reduce $Cu(II)$ to Cu(I) but bromide cannot.

$4I^- + 2Cu^{2+} \rightarrow Cu_2I_2 + I_2$

Hence, among the hydrohalic compounds, hydroiodic acid is the best reductant. The reducing power of hydrohalic acids is $HF < HCl < HBr < HI$.

#423692

Topic: Oxidation Number

Passage

Consider the elements Cs, Ne, I and F.

Identify the element that exhibits only negative oxidation state.

Solution

Fluorine is the most electronegative element in the periodic table. It exhibits only negative oxidation state of -1.

#423693

Topic: Oxidation Number

Passage

Consider the elements Cs, Ne, I and F.

Identify the element that exhibits only postive oxidation state.

Solution

Cs is the most electropositive element in the periodic table. It is an alkali metal and exhibits only positive oxidation state of +1.

#423694

Topic: Oxidation Number

Passage

Consider the elements Cs, Ne, I and F.

Identify the element that exhibits both positive and negative oxidation states.

Solution

Iodine exhibits both positive and negative oxidation states. It exhibits oxidation states -1, +1, +3, +5 and +7.

#423695

Topic: Oxidation Number

Passage

Consider the elements Cs, Ne, I and F.

Identify the element which exhibits neither the negative nor does the positive oxidation state.

Solution

The element which exhibits neither the negative nor does the positive oxidation state is Ne. It is a noble gas with oxidation state of zero.

#423701

Topic: Balance redox reactions

In Ostwald's process for the manufacture of nitric acid, the first step involves the oxidation of ammonia gas by oxygen gas to give nitric oxide gas and steam. What is the maximum weight of nitric oxide that can be obtained starting only with 10.00 g. of ammonia and 20.00 g of oxygen ?

Solution

The balanced chemical equation is:

$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$

The molar masses of ammonia and oxygen are 17 g/mol and 32 g/mol respectively.

5 moles (160 g) of oxygen reacts with 4 moles (68 g) of ammonia.

20 g of oxygen will react with $\frac{68 \times 20}{160} = 8.5$ g of ammonia. Hence, oxygen is the limiting regent.

The molar mass of NO is 30 g/mol.

5 moles (160 g) of oxygen will produce 4 moles (120 g) of $NO$$.

20 g of oxygen will produce $\frac{120 \times 20}{160} = 15g$ of NO.

#423705

Topic: Electrode potential

Increasing strength of oxidising agent ↑ | Increasing strength of reducing agent ↓

Reaction (Oxidised form + ne^-	→ Reduced form)	$E^\ominus$ / V
$F_2(g) + 2e^-$	$\rightarrow 2F^-$	2.87
$Co^{3+} + e^-$	$\rightarrow Co^{2+}$	1.81
$H_2O_2 + 2H^+ + 2e^-$	$\rightarrow 2H_2O$	1.78
$MnO_4^- + 8H^+ + 5e^-$	$\rightarrow Mn^{2+} + 4H_2O$	1.51
$Au^{3+} + 3e^-$	$\rightarrow Au(s)$	1.40
$Cl_2(g) + 2e^-$	$\rightarrow 2Cl^-$	1.36
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	$\rightarrow 2Cr^{3+} + 7H_2O$	1.33
$O_2(g) + 4H^+ + 4e^-$	$\rightarrow 2H_2O$	1.23
$MnO_2(s) + 4H^+ + 2e^-$	$\rightarrow Mn^{2+} + 2H_2O$	1.23
$Br_2 + 2e^-$	$\rightarrow 2Br^-$	1.09
$NO_3^- + 4H^+ + 3e^-$	$\rightarrow NO(g) + 2H_2O$	0.97
$2Hg^{2+} + 2e^-$	$\rightarrow Hg_2^{2+}$	0.92
$Ag^+ + e^-$	$\rightarrow Ag(s)$	0.80
$Fe^{3+} + e^-$	$\rightarrow Fe^{2+}$	0.77
$O_2(g) + 2H^+ + 2e^-$	$\rightarrow H_2O_2$	0.68
$I_2(s) + 2e^-$	$\rightarrow 2I^-$	0.54
$Cu^+ + e^-$	$\rightarrow Cu(s)$	0.52
$Cu^{2+} + 2e^-$	$\rightarrow Cu(s)$	0.34
$AgCl(s) + e^-$	$\rightarrow Ag(s) + Cl^-$	0.22
$AgBr(s) + e^-$	$\rightarrow Ag(s) + Br^-$	0.10
$2H^+ + 2e^-$	**$\rightarrow H_2(g)$**	**0.00**
$Pb^{2+} + 2e^-$	$\rightarrow Pb(s)$	–0.13
$Sn^{2+} + 2e^-$	$\rightarrow Sn(s)$	–0.14
$Ni^{2+} + 2e^-$	$\rightarrow Ni(s)$	–0.25
$Fe^{2+} + 2e^-$	$\rightarrow Fe(s)$	–0.44
$Cr^{3+} + 3e^-$	$\rightarrow Cr(s)$	–0.74
$Zn^{2+} + 2e^-$	$\rightarrow Zn(s)$	–0.76
$2H_2O + 2e^-$	$\rightarrow H_2(g) + 2OH^-$	–0.83
$Al^{3+} + 3e^-$	$\rightarrow Al(s)$	–1.66
$Mg^{2+} + 2e^-$	$\rightarrow Mg(s)$	–2.36
$Na^+ + e^-$	$\rightarrow Na(s)$	–2.71
$Ca^{2+} + 2e^-$	$\rightarrow Ca(s)$	–2.87
$K^+ + e^-$	$\rightarrow K(s)$	–2.93
$Li^+ + e^-$	$\rightarrow Li(s)$	–3.05

Using the standard electrode potentials given in the table, predict if the reaction between the following is possible.

$Fe^{3+}(aq)$ and $I^-(aq)$

Solution

The oxidation half reaction is $2I^-(aq) \rightarrow I_2(s) + 2e^-; E^0 = -0.54V$.

The reduction half reaction is $[Fe^{3+}(aq) + e^- \rightarrow] \times 2; E^0 = +0.77 \times 2\ V = +0.77V$.

The net cell reaction is $2Fe^{3+}(aq) + 2I^-(aq) \rightarrow 2Fe^{2+}(aq) + I_2(s); E^0 = +23V$.

Since, the cell potential is positive, the reaction is feasible.

#423706

Topic: Electrode potential

Reaction (Oxidised form + ne^-	$\rightarrow$ Reduced form)	$E^\ominus$ / V
$F_2(g) + 2e^-$	$\rightarrow 2F^-$	2.87
$Co^{3+} + e^-$	$\rightarrow Co^{2+}$	1.81
$H_2O_2 + 2H^+ + 2e^-$	$\rightarrow 2H_2O$	1.78
$MnO_4^- + 8H^+ + 5e^-$	$\rightarrow Mn^{2+} + 4H_2O$	1.51
$Au^{3+} + 3e^-$	$\rightarrow Au(s)$	1.40
$Cl_2(g) + 2e^-$	$\rightarrow 2Cl^-$	1.36
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	$\rightarrow 2Cr^{3+} + 7H_2O$	1.33
$O_2(g) + 4H^+ + 4e^-$	$\rightarrow 2H_2O$	1.23
$MnO_2(s) + 4H^+ + 2e^-$	$\rightarrow Mn^{2+} + 2H_2O$	1.23
$Br_2 + 2e^-$	$\rightarrow 2Br^-$	1.09
$NO_3^- + 4H^+ + 3e^-$	$\rightarrow NO(g) + 2H_2O$	0.97
$2Hg^{2+} + 2e^-$	$\rightarrow Hg_2^{2+}$	0.92
$Ag^+ + e^-$	$\rightarrow Ag(s)$	0.80
$Fe^{3+} + e^-$	$\rightarrow Fe^{2+}$	0.77
$O_2(g) + 2H^+ + 2e^-$	$\rightarrow H_2O_2$	0.68
$I_2(s) + 2e^-$	$\rightarrow 2I^-$	0.54
$Cu^+ + e^-$	$\rightarrow Cu(s)$	0.52
$Cu^{2+} + 2e^-$	$\rightarrow Cu(s)$	0.34
$AgCl(s) + e^-$	$\rightarrow Ag(s) + Cl^-$	0.22
$AgBr(s) + e^-$	$\rightarrow Ag(s) + Br^-$	0.10
$2H^+ + 2e^-$	**$\rightarrow H_2(g)$**	**0.00**
$Pb^{2+} + 2e^-$	$\rightarrow Pb(s)$	–0.13
$Sn^{2+} + 2e^-$	$\rightarrow Sn(s)$	–0.14
$Ni^{2+} + 2e^-$	$\rightarrow Ni(s)$	–0.25
$Fe^{2+} + 2e^-$	$\rightarrow Fe(s)$	–0.44
$Cr^{3+} + 3e^-$	$\rightarrow Cr(s)$	–0.74
$Zn^{2+} + 2e^-$	$\rightarrow Zn(s)$	–0.76
$2H_2O + 2e^-$	$\rightarrow H_2(g) + 2OH^-$	–0.83
$Al^{3+} + 3e^-$	$\rightarrow Al(s)$	–1.66
$Mg^{2+} + 2e^-$	$\rightarrow Mg(s)$	–2.36
$Na^+ + e^-$	$\rightarrow Na(s)$	–2.71
$Ca^{2+} + 2e^-$	$\rightarrow Ca(s)$	–2.87
$K^+ + e^-$	$\rightarrow K(s)$	–2.93
$Li^+ + e^-$	$\rightarrow Li(s)$	–3.05

Increasing strength of oxidising agent

Increasing strength of reducing agent

Using the standard electrode potentials given in the table, predict if the reaction between the following is possible.

Ag^+(aq) and $Cu(s)$

Solution

Oxidation half reaction is $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-; E^0 = -0.34V$.

Reduction half reaction is $[Ag^+(aq) + e^- \rightarrow Ag(s)] \times 2; E^0 = +0.80V$.

The net cell reaction is $2Ag^+(aq) + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}; E^0 = +0.46V$.

Since, the cell potential is positive, the reaction is feasible.

#423708

Topic: Electrode potential

Reaction (Oxidised form + ne^-	$\rightarrow$ Reduced form)	$E^\ominus$ / V
$F_2(g) + 2e^-$	$\rightarrow 2F^-$	2.87
$Co^{3+} + e^-$	$\rightarrow Co^{2+}$	1.81
$H_2O_2 + 2H^+ + 2e^-$	$\rightarrow 2H_2O$	1.78
$MnO_4^- + 8H^+ + 5e^-$	$\rightarrow Mn^{2+} + 4H_2O$	1.51
$Au^{3+} + 3e^-$	$\rightarrow Au(s)$	1.40
$Cl_2(g) + 2e^-$	$\rightarrow 2Cl^-$	1.36
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	$\rightarrow 2Cr^{3+} + 7H_2O$	1.33
$O_2(g) + 4H^+ + 4e^-$	$\rightarrow 2H_2O$	1.23
$MnO_2(s) + 4H^+ + 2e^-$	$\rightarrow Mn^{2+} + 2H_2O$	1.23
$Br_2 + 2e^-$	$\rightarrow 2Br^-$	1.09
$NO_3^- + 4H^+ + 3e^-$	$\rightarrow NO(g) + 2H_2O$	0.97
$2Hg^{2+} + 2e^-$	$\rightarrow Hg_2^{2+}$	0.92
$Ag^+ + e^-$	$\rightarrow Ag(s)$	0.80
$Fe^{3+} + e^-$	$\rightarrow Fe^{2+}$	0.77
$O_2(g) + 2H^+ + 2e^-$	$\rightarrow H_2O_2$	0.68
$I_2(s) + 2e^-$	$\rightarrow 2I^-$	0.54
$Cu^+ + e^-$	$\rightarrow Cu(s)$	0.52
$Cu^{2+} + 2e^-$	$\rightarrow Cu(s)$	0.34
$AgCl(s) + e^-$	$\rightarrow Ag(s) + Cl^-$	0.22
$AgBr(s) + e^-$	$\rightarrow Ag(s) + Br^-$	0.10
$2H^+ + 2e^-$	**$\rightarrow H_2(g)$**	**0.00**
$Pb^{2+} + 2e^-$	$\rightarrow Pb(s)$	–0.13
$Sn^{2+} + 2e^-$	$\rightarrow Sn(s)$	–0.14
$Ni^{2+} + 2e^-$	$\rightarrow Ni(s)$	–0.25
$Fe^{2+} + 2e^-$	$\rightarrow Fe(s)$	–0.44
$Cr^{3+} + 3e^-$	$\rightarrow Cr(s)$	–0.74
$Zn^{2+} + 2e^-$	$\rightarrow Zn(s)$	–0.76
$2H_2O + 2e^-$	$\rightarrow H_2(g) + 2OH^-$	–0.83
$Al^{3+} + 3e^-$	$\rightarrow Al(s)$	–1.66
$Mg^{2+} + 2e^-$	$\rightarrow Mg(s)$	–2.36
$Na^+ + e^-$	$\rightarrow Na(s)$	–2.71
$Ca^{2+} + 2e^-$	$\rightarrow Ca(s)$	–2.87
$K^+ + e^-$	$\rightarrow K(s)$	–2.93
$Li^+ + e^-$	$\rightarrow Li(s)$	–3.05

Increasing strength of oxidising agent

Increasing strength of reducing agent

Using the standard electrode potentials given in the table, predict if the reaction between the following is possible.

Fe^{3+}(aq) and $Cu(s)$

Solution

Oxidation half reaction is $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-; E^0 = -0.34V$.

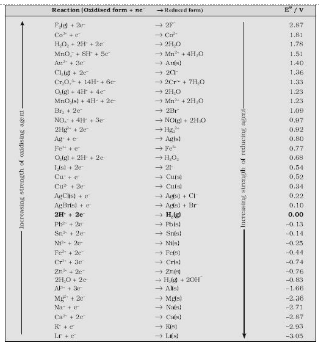
Reduction half reaction is $[Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)] \times 2; E^0 = +0.77V$.

The net cell reaction is $2Fe^{3+}(aq) + Cu(s) \rightarrow 2Fe^{2+}(aq) + Cu^{2+}; E^0 = +0.43V$.

Since, the cell potential is positive, the reaction is feasible.

#423710

Topic: Types of redox reactions



Reaction (Oxidised form + ne^-	$\rightarrow$ Reduced form)	$E^\ominus$ / V
$F_2(g) + 2e^-$	$\rightarrow 2F^-$	2.87
$Co^{3+} + e^-$	$\rightarrow Co^{2+}$	1.81
$H_2O_2 + 2H^+ + 2e^-$	$\rightarrow 2H_2O$	1.78
$MnO_4^- + 8H^+ + 5e^-$	$\rightarrow Mn^{2+} + 4H_2O$	1.51
$Au^{3+} + 3e^-$	$\rightarrow Au(s)$	1.40
$Cl_2(g) + 2e^-$	$\rightarrow 2Cl^-$	1.36
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	$\rightarrow 2Cr^{3+} + 7H_2O$	1.33
$O_2(g) + 4H^+ + 4e^-$	$\rightarrow 2H_2O$	1.23
$MnO_2(s) + 4H^+ + 2e^-$	$\rightarrow Mn^{2+} + 2H_2O$	1.23
$Br_2 + 2e^-$	$\rightarrow 2Br^-$	1.09
$NO_3^- + 4H^+ + 3e^-$	$\rightarrow NO(g) + 2H_2O$	0.97
$2Hg^{2+} + 2e^-$	$\rightarrow Hg_2^{2+}$	0.92
$Ag^+ + e^-$	$\rightarrow Ag(s)$	0.80
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$O_2(g) + 2H^+ + 2e^-$	$\rightarrow H_2O_2$	0.68
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$Cu^+ + e^-$	$\rightarrow Cu(s)$	0.52
$Cu^{2+} + 2e^-$	$\rightarrow Cu(s)$	0.34
$AgCl(s) + e^-$	$\rightarrow Ag(s) + Cl^-$	0.22
$AgBr(s) + e^-$	$\rightarrow Ag(s) + Br^-$	0.10
$2H^+ + 2e^-$	**$\rightarrow H_2(g)$**	**0.00**
$Pb^{2+} + 2e^-$	$\rightarrow Pb(s)$	–0.13
$Sn^{2+} + 2e^-$	$\rightarrow Sn(s)$	–0.14
$Ni^{2+} + 2e^-$	$\rightarrow Ni(s)$	–0.25
$Fe^{2+} + 2e^-$	$\rightarrow Fe(s)$	–0.44
$Cr^{3+} + 3e^-$	$\rightarrow Cr(s)$	–0.74
$Zn^{2+} + 2e^-$	$\rightarrow Zn(s)$	–0.76
$2H_2O + 2e^-$	$\rightarrow H_2(g) + 2OH^-$	–0.83
$Al^{3+} + 3e^-$	$\rightarrow Al(s)$	–1.66
$Mg^{2+} + 2e^-$	$\rightarrow Mg(s)$	–2.36
$Na^+ + e^-$	$\rightarrow Na(s)$	–2.71
$Ca^{2+} + 2e^-$	$\rightarrow Ca(s)$	–2.87
$K^+ + e^-$	$\rightarrow K(s)$	–2.93
$Li^+ + e^-$	$\rightarrow Li(s)$	–3.05

Increasing strength of oxidising agent

Increasing strength of reducing agent

Using the standard electrode potentials given in the Table, predict if the reaction between the following is feasible.

$Ag(s)$ and $Fe^{3+}(aq)$

Solution

Oxidation half reaction is $Ag(s) \rightarrow Ag^+(aq) + e^-; E^0 = -0.80V$.

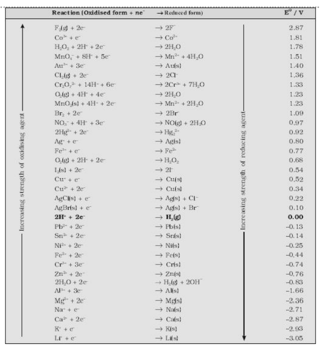
Reduction half reaction is $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq); E^0 = +0.77V$.

The net cell reaction is $2Fe^{3+}(aq) + Ag(s) \rightarrow 2Fe^{2+}(aq) + Ag^+; E^0 = -0.03V$.

Since, the cell potential is negative, the reaction is not feasible.

#423712

Topic: Electrode potential



Reaction (Oxidised form + ne^-	$\rightarrow$ Reduced form)	$E^\ominus$ / V
$F_2(g) + 2e^-$	$\rightarrow 2F^-$	2.87
$Co^{3+} + e^-$	$\rightarrow Co^{2+}$	1.81
$H_2O_2 + 2H^+ + 2e^-$	$\rightarrow 2H_2O$	1.78
$MnO_4^- + 8H^+ + 5e^-$	$\rightarrow Mn^{2+} + 4H_2O$	1.51
$Au^{3+} + 3e^-$	$\rightarrow Au(s)$	1.40
$Cl_2(g) + 2e^-$	$\rightarrow 2Cl^-$	1.36
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	$\rightarrow 2Cr^{3+} + 7H_2O$	1.33
$O_2(g) + 4H^+ + 4e^-$	$\rightarrow 2H_2O$	1.23
$MnO_2(s) + 4H^+ + 2e^-$	$\rightarrow Mn^{2+} + 2H_2O$	1.23
$Br_2 + 2e^-$	$\rightarrow 2Br^-$	1.09
$NO_3^- + 4H^+ + 3e^-$	$\rightarrow NO(g) + 2H_2O$	0.97
$2Hg^{2+} + 2e^-$	$\rightarrow Hg_2^{2+}$	0.92
$Ag^+ + e^-$	$\rightarrow Ag(s)$	0.80
$Fe^{3+} + e^-$	$\rightarrow Fe^{2+}$	0.77
$O_2(g) + 2H^+ + 2e^-$	$\rightarrow H_2O_2$	0.68
$I_2(s) + 2e^-$	$\rightarrow 2I^-$	0.54
$Cu^+ + e^-$	$\rightarrow Cu(s)$	0.52
$Cu^{2+} + 2e^-$	$\rightarrow Cu(s)$	0.34
$AgCl(s) + e^-$	$\rightarrow Ag(s) + Cl^-$	0.22
$AgBr(s) + e^-$	$\rightarrow Ag(s) + Br^-$	0.10
$2H^+ + 2e^-$	**$\rightarrow H_2(g)$**	**0.00**
$Pb^{2+} + 2e^-$	$\rightarrow Pb(s)$	–0.13
$Sn^{2+} + 2e^-$	$\rightarrow Sn(s)$	–0.14
$Ni^{2+} + 2e^-$	$\rightarrow Ni(s)$	–0.25
$Fe^{2+} + 2e^-$	$\rightarrow Fe(s)$	–0.44
$Cr^{3+} + 3e^-$	$\rightarrow Cr(s)$	–0.74
$Zn^{2+} + 2e^-$	$\rightarrow Zn(s)$	–0.76
$2H_2O + 2e^-$	$\rightarrow H_2(g) + 2OH^-$	–0.83
$Al^{3+} + 3e^-$	$\rightarrow Al(s)$	–1.66
$Mg^{2+} + 2e^-$	$\rightarrow Mg(s)$	–2.36
$Na^+ + e^-$	$\rightarrow Na(s)$	–2.71
$Ca^{2+} + 2e^-$	$\rightarrow Ca(s)$	–2.87
$K^+ + e^-$	$\rightarrow K(s)$	–2.93
$Li^+ + e^-$	$\rightarrow Li(s)$	–3.05

Increasing strength of oxidising agent

Increasing strength of reducing agent

Using the standard electrode potentials given in the table, predict if the reaction between the following is possible.

$Br_2(aq)$ and $Fe^{2+}(aq)$

Solution

Oxidation half reaction is $[Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-]; E^0 = -0.77V$.

Reduction half reaction is $Br_2(aq) + 2e^- \rightarrow 2Br^-(aq); E^0 = +1.09V$.

The net cell reaction is $Br_2(aq) + 2Fe^{2+}(aq) \rightarrow 2Br^-(aq) + 2Fe^{3+}(aq); E^0 = -0.32V$.

Since, the cell potential is negative, the reaction is not feasible.

#423716

Topic: Types of redox reactions

Passage

Predict the products of electrolysis in each of the following.

An aqueous solution of $AgNO_3$ with silver electrodes.

Solution

Reaction in solution;

$AgNO_3 \rightarrow Ag^+ + NO_3^-$

$H_2O \rightarrow H^+ + OH^-$

Reaction at cathode;

$Ag^+ + e^- \rightarrow Ag$

Reaction at anode;

$Ag(s) + NO_3^- \rightarrow AgNO_3(aq) + e^-$

In aqueous solution, silver nitrate ionizes to silver ions and nitrate ions.

At cathode, either silver ions or water molecules can be reduced.

Since, silver ion has higher reduction potential than water, silver ions are reduced at cathode.

Similarly, at anode, either silver metal of water molecules can be oxidized. Since oxidation potential of silver is higher than that of water molecules, silver is oxidized.

#423717

Topic: Types of redox reactions

Passage

Predict the products of electrolysis in each of the following.

An aqueous solution $AgNO_3$ with platinum electrodes.

Solution

The oxidation of Pt is not possible. Water is oxidized at anode which liberates oxygen. Silver ions are reduced at cathode and are deposited.

Reaction in solution;

$AgNO_3 \rightarrow Ag^+ + NO_3^-$

$H_2O \rightarrow H^+ + OH^-$

Reaction at cathode;

$Ag^+ + e^- \rightarrow Ag$

Reaction at anode;

Due to platinum electrode, self ionization of water will take place.

$H_2O \rightarrow 2H^+ + \frac{1}{2}O_2 + 2e^-$

Hence, silver will deposit at cathode and oxygen gas will generate at anode.

#423718

Topic: Types of redox reactions

Passage

Predict the products of electrolysis in each of the following.

A dilute solution of H_2SO_4 with platinum electrodes.

Solution

The dissociation of sulphuric acid gives protons and sulphate ions.

At cathode, either hydrogen ions or water molecules can be reduced. Since protons have higher reduction potential than water, hydrogen ions are reduced to hydrogen gas.

At anode, either sulphate ions or water molecule can get oxidized. Since, during oxidation of sulphate ions, more bonds are broken than oxidation of water molecules, sulphate ions have lower oxidation potential than water. Hence, water is oxidized at the anode to liberate oxygen molecules.

$H_2SO_4 \rightarrow 2H^+ + SO_4^-$

$H_2O \rightarrow H^+ + OH^-$

Reaction at cathode;

$2H^+ + 2e^- \rightarrow H_2$

Reaction at anode;

Due to platinum electrode, self ionization of water will take place.

$H_2O \rightarrow 2H^+ + \frac{1}{2}O_2 + 2e^-$

Hence, hydrogen gas will generate at cathode and oxygen gas will generate at anode.

#423719

Topic: Types of redox reactions

Passage

Predict the products of electrolysis in each of the following.

An aqueous solution of $CuCl_2$ with platinum electrodes.

Solution

When an aqueous solution of $CuCl_2$ is electrolyzed with platinum electrodes, chlorine is obtained at anode and Cu is deposited at cathode.

$2Cl^- \rightarrow Cl_2 + 2e^-$

$Cu^{2+} + 2e^- \rightarrow Cu$

#423720

Topic: Electrode potential

Arrange the following metals in the order in which they displace each other from the solution of their salts.

Al, Cu, Fe, Mg and Zn.

Solution

A metal having stronger reducing power displaces another metal having weaker reducing power from its salt solution.

The increasing order of reducing power is $Cu < Fe < Zn < Al < Mg$.

Thus, Mg can displace Al from its salt solution but Al cannot displace Mg.

The order in which the metals displace each other from their salt solutions is $Mg > Al > Zn > Fe > Cu$.

#423721

Topic: Electrode potential

Given the standard electrode potentials of some metals.

$K^+/K = 2.93V$,

$Ag^+/Ag = 0.80V$,

$Hg^{2+}/Hg = 0.79V$,

$Mg^{2+}/Mg = 2.37V$ and

$Cr^{3+}/Cr = 0.74V$

Arrange these metals in their increasing order of reducing power.

Solution

Lower reduction potential corresponds to higher reducing power. The increasing order of the standard reduction potentials is

$K^+ | K < Mg^{2+} | Mg < Cr^{3+} | Cr < Hg^{2+} | Hg < Ag^+ | Ag$.

The increasing order of the reducing power is $Ag < Hg < Cr < Mg < K$.

#423723

Topic: Electrode potential

Passage

Depict the galvanic cell in which the reaction $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$ takes place. Further show:

which of the electrode is negatively charged?

Solution

For the given redox reaction, the galvanic cell is :

$Zn \mid Zn^{2+}(aq) \mid \mid Ag^+(aq) \mid Ag$

Zinc electrode is negatively charged as Zn is oxidized to Zn^{2+}. The electrons released during oxidation accumulate on this electrode.

#423725

Topic: Electrode potential

Passage

Depict the galvanic cell in which the reaction $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$ takes place. Further show:

individual reaction at each electrode.

Solution

For the given redox reaction, the galvanic cell is

$Zn \mid Zn^{2+}(aq) \mid \mid Ag^+(aq) \mid Ag$

At Zn electrode, Zn is oxidized to $Zn(II)$ ions.

At Ag electrode, $Ag(I)$ is reduced to Ag.

#464730

Topic: Types of redox reactions

Food cans are coated with tin and not with zinc because:

A zinc is costlier than tin

B zinc has a higher melting point than tin

$\boxed{\text{C}}$ zinc is more reactive than tin

D zinc is less reactive than tin

Solution

Food cans are coated with tin and not with zinc because zinc is above the tin in reactivity series means more reactive than tin and can react with food elements preserved in it.

#423521
Topic: Oxidation Number

Passage

Justify that the following reactions are redox reactions:

$Fe_2O_3(s)\ +\ 3CO(g)\ \rightarrow 2Fe(s)\ +\ 3CO_2(g)$

Solution

The oxidation number of iron decreases from +3 to 0. The oxidation number of C increases from +2 to +4. Hence, Fe_2O_3 is reduced and CO is oxidized. Hence, it is a redox reaction.

#423532
Topic: Oxidation Number

Passage

Justify that the following reactions are redox reactions:

$4BCl_3(g)\ +\ 3LiAlH_4(g)\ \rightarrow 2B_2H_6(g)\ +\ 3LiCl(s)\ +\ 3\ AlCl_3(s)$

Solution

The oxidation number of B decreases from +3 to -3. The oxidation number of hydrogen increases from -1 to +1. Hence, BCl_3 is reduced and $LiAlH_4$ is oxidized.

Hence, it is a redox reaction.

#423537
Topic: Oxidation Number

Passage

Justify that the following reactions are redox reactions:

$2K(s)\ +\ F_2(g)\ \rightarrow 2K^+F^-(s)$

Solution

The oxidation number of K increases from 0 to +1 and the oxidation number of F_2 decreases from 0 to -1. Hence, K is oxidized and F_2 is reduced. Hence, it is redox reaction.

#423542
Topic: Oxidation and reduction - classical concept

Consider the reaction of water with F_2 and suggest in terms of oxidation and reduction which species are oxidised/reduced.

Solution

The balanced chemical equations are given below.

$2F_2 + 2H_2O \rightarrow O_2 + 4H^+ + 4F^-$

$3F_2 + 3H_2O \rightarrow O_3 + 6H^+ + 6F^-$

Thus, water is a reductant and itself is oxidized to oxygen or ozone. Fluorine is an oxidant and itself is reduced to fluoride ion.

#423543
Topic: Oxidation Number

Passage

Justify that the following reactions are redox reactions:

$4NH_3(g)\ +\ 5O_2(g)\ \rightarrow 4NO(g)\ +\ 6H_2(s)$

Solution

The oxidation number of nitrogen increases from -3 to +2. The oxidation number of oxygen decreases from 0 to -2. Ammonia is oxidized and oxygen is reduced. Thus, it is a redox reaction.

#423584

Topic: Balance redox reactions

Consider the reactions:

Why it is more appropriate to write these reactions as: $6\,CO_2(g)\,+\,12H_2O(l)\,\rightarrow\,C_6H_{12}O_6(aq)\,+\,6O_2(g)+6H_2O(l)$

Also suggest a technique to investigate the path of the above (a) and (b) redox reactions.

$6\,CO_2(g)\,+\,6H_2O(l)\,\rightarrow\,C_6H_{12}O_6(aq)\,+\,6O_2(g)$

Solution

The given reaction is

$6\,CO_2(g)\,+\,6H_2O(l)\,\rightarrow\,C_6H_{12}O_6(aq)\,+\,6O_2(g)$

It is more appropriate to write this reaction as

$6\,CO_2(g)\,+\,12H_2O(l)\,\rightarrow\,C_6H_{12}O_6(aq)\,+\,6O_2(g)+6H_2O(l)$

This is because water is produced during photosynthesis and water must be shown on product side. To investigate the path, H_2O^{18} is used instead of H_2O^{16}. Thus, instead of using normal O atom in water molecule, we use radioactively labelled O atom in water molecule. By determining the intermediates and products containing labelled oxygen, we can investigate above paths.

#423618

Topic: Balance redox reactions

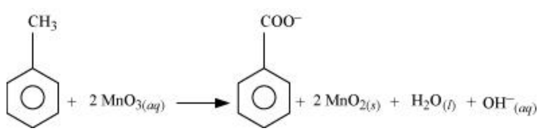
Though alkaline potassium permanganate and acidic potassium permanganate both are used as oxidants, yet in the manufacture of benzoic acid from toluene we use alcoholic potassium permanganate as an oxidant. Why? Write the balanced redox equation for the reaction.

Solution

Toluene is oxidized to benzoic acid by using alcoholic potassium permanganate.

In neutral medium hydroxide ions are produced. This reduces cost of adding acid or base.

Alcohol and $KMnO_4$ are both polar and homogeneous to each other. Toluene and alcohol are homogeneous as they are organic. Thus, in alcohol the rate of reaction between toluene and $KMnO_4$ is higher. The balanced redox reaction is as shown.



#423621

Topic: Types of redox reactions

How do you count for the following observations?

When concentrated sulphuric acid is added to an inorganic mixture containing chloride, we get colourless pungent smell gas HCl, but if the mixture contains bromide then we get red vapour of bromine. Why ?

Solution

When concentrated sulphuric acid is added to an inorganic mixture containing chloride, we get colourless pungent smelling gas HCl, but if the mixture contains bromide then we get red vapour of bromine. This is because HBr can be formed only if dil H_2SO_4 is used. Note: concentrated sulphuric acid converts inorganic chloride to HCl. But concentrated sulphuric acid converts inorganic bromide to bromine. Dilute sulphuric acid will convert inorganic bromide to HBr.

#423622

Topic: Types of redox reactions

When concentrated sulphuric acid is added to an inorganic mixture containing chloride, we get colorless pungent smell gas HCl, but if the mixture contains bromide then we get red vapour of bromine. Why ?

Solution

When concentrated sulphuric acid is added to an inorganic mixture containing chloride, we get colourless pungent smelling gas HCl, but if the mixture contains bromide then we get red vapour of bromine. This is because HBr can be formed only if dil H_2SO_4 is used.

Note: concentrated sulphuric acid converts inorganic chloride to HCl.

But concentrated sulphuric acid converts inorganic bromide to bromine.

Dilute sulphuric acid will convert inorganic bromide to HBr.

#423632

Topic: Oxidation Number

Passage

Identify the substance oxidised, reduced, oxidising agent and reducing agent for each of the following reactions.

$2AgBr(s)\ +\ C_6H_6O_2(aq)\ \rightarrow\ 2Ag(s)\ +\ 2H\,Br(aq)\ +\ C_6H_4O_2(aq)$

Solution

Ag^+ is reduced and acts as oxidizing agent.

$C_6H_6O_2$ is oxidized and acts as reducing agent.

#423636

Topic: Oxidation Number

Passage

Identify the substance oxidised, reduced, oxidising agent and reducing agent for each of the following reactions.

$HCHO(l)\ +\ 2[Ag\,(NH_3)_2]^+\,(aq)\ +\ 3OH^-(aq)\ \rightarrow\ 2Ag(s)\ +\ HCOO^-(aq)\ +\ 4NH_3(aq)\ +\ 2H_2O(l)$

Solution

$HCHO$ is oxidized and $[Ag(NH_3)_2]^+$ is reduced.

$[Ag(NH_3)_2]^+$ is the oxidizing agent and $HCHO$ is reducing agent.

#423659

Topic: Oxidation Number

Passage

Identify the substance oxidised, reduced, oxidising agent and reducing agent for each of the following reactions.

$HCHO\,(1) + 2Cu^{2+}(aq) + 5OH^-(aq) \rightarrow Cu_2O(s) + HCOO^-(aq) + 3H_2O(l)$

Solution

$HCHO$ is oxidized and Cu^{2+} is reduced.

Cu^{2+} is the oxidizing agent and $HCHO$ is the reducing agent.

#423661

Topic: Oxidation Number

Passage

Identify the substance oxidised, reduced, oxidising agent and reducing agent for each of the following reactions.

$Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$

Solution

Pb is oxidized ad PbO_2 is reduced.

PbO_2 is oxidizing agent and Pb is reducing agent.

#423666

Topic: Oxidation and reduction - electron transfer concept

$XeO_6^{4-}(aq)+2F^-(aq)+6H^+(aq)\rightarrow XeO_3(g)+3H_2O(l)$

What conclusion about the compound Na_4XeO_6(of which XeO_6^{4-} is a part) can be drawn from the reaction?

Solution

XeO_6^{4-} oxidizes F^- and F^- reduces XeO_6^{4-}.

Hence, the given reaction occurs.

The oxidation number of Xe decreases from +8 to +6. The oxidation number of F increases from -1 to 0.

Thus, Na_4XeO_6 is a stronger oxidsing agent than F^-.

#423668

Topic: Oxidation and reduction - electron transfer concept

From the following reactions, determine if Ag^+ is a stronger oxidizing agent than Cu^{2+}.

(a) $H_3PO_2(aq)+4AgNO_3(aq)+2H_2O(l)\rightarrow H_3PO(aq)+4Ag(s)+4HNO_3(aq)$

(b) $H_3PO_2(aq)+2CuSO_4(aq)+2H_2O(l)\rightarrow H_3PO_4(aq)+2Cu(s)+H_2SO_4(aq)$

(c) $C_6H_5CHO(l)+2[Ag(NH_3)_2]^+(aq)+3OH^-(aq)\rightarrow C_6H_5COO^-(aq)+2Ag(s)+4NH_3(aq)+2H_2O(l)$

(d) $C_6H_5CHO(l)+2Cu^{2+}(aq)+5OH^-(aq)\rightarrow$No Change observed.

Solution

In the reactions (a) and (b), Ag^+ and Cu^{2+} are oxidizing agents. In the reaction (c) Ag^+ oxidizes benzaldehyde to benzoate ion. In reaction (d) Cu^{2+} cannot oxidize benzaldehyde. Hence, Ag^+ is a stronger oxidizing agent than Cu^{2+} .

#423669

Topic: Oxidation and reduction - electron transfer concept

From the following reactions, determine if Ag^+ is a stronger oxidizing agent than Cu^{2+}

(a) $H_3PO_2(aq)+4AgNO_3(aq)+2H_2O(l)\rightarrow H_3PO(aq)+4Ag(s)+4HNO_3(aq)$

(b) $H_3PO_2(aq)+2CuSO_4(aq)+2H_2O(l)\rightarrow H_3PO_4(aq)+2Cu(s)+H_2SO_4(aq)$

(c) $C_6H_5CHO(l)+2[Ag(NH_3)_2]^+(aq)+3OH^-(aq)\rightarrow C_6H_5COO^-(aq)+2Ag(s)+4NH_3(aq)+2H_2O(l)$

(d) $C_6H_5CHO(l)+2Cu^{2+}(aq)+5OH^-(aq)\rightarrow$No Change observed.

Solution

In the reactions (a) and (b), Ag^+ and Cu^{2+} are oxidizing agents. In the reaction (c) Ag^+

oxidizes benzaldehyde to benzoate ion. In reaction (d) Cu^{2+} cannot oxidize benzaldehyde. Hence, Ag^+ is a stronger oxidizing agent than Cu^{2+}

#423674

Topic: Balance redox reactions

Balance the following redox reactions by ion electron method.

$MnO_4^-(aq)+I^-(aq)\rightarrow MnO_2(s)+I_2(s)$(in basis medium)

Solution

The unbalanced chemical equation is

$MnO_4^-(aq) + I^-(aq) \rightarrow MnO_2(s) + I_2(s)$

The oxidation half reaction is

$I^-(aq) \rightarrow I_2(s)$

The reduction half reaction is

$MnO_4^-(aq) \rightarrow MnO_2(aq)$

Balance I atoms and charges in the oxidation half reaction.

$2I^-(aq) \rightarrow I_2(s) + 2e^-$

In the reduction half reaction, the oxidation number of Mn changes from +7 to +4. Hence, add 3 electrons to reactant side of the reaction.

$MnO_4^-(aq) + 3e^- \rightarrow MnO_2(aq)$

Balance charge in the reduction half reaction by adding 4 hydroxide

ions to product side.

$MnO_4^-(aq) + 3e^- \rightarrow MnO_2(aq) + 4OH^-$

To balance O atoms, add 2 water molecules to reactant side.

$MnO_4^-(aq) + 3e^- + 2H_2O \rightarrow MnO_2(aq) + 4OH^-$

To equalize the number of electrons, multiply the oxidation half reaction by 3 and multiply the reduction half reaction by 2.

$6I^-(aq) \rightarrow 3I_2(s) + 6e^-$

$2MnO_4^-(aq) + 6e^- + 4H_2O \rightarrow 2MnO_2(aq) + 8OH^-$

Add two half cell reactions to obtain the balanced equation.

$2MnO_4^-(aq) + 6I^-(aq) + 4H_2O_2(l) \rightarrow 2MnO_2(s) + 3I_2(s) + 8OH^-$

#423675

Topic: Balance redox reactions

Passage

Balance the following redox reactions by ion electron method.

$MnO_4^-(aq) + SO_2(g) \rightarrow Mn^{2+}(aq) + HSO_4^-(aq)$(in acidic solution)

Solution

The unbalanced chemical equation is:

$MnO_4^-(aq) + SO_2(g) \rightarrow Mn^{2+}(aq) + HSO_4^-(aq)$

The oxidation half reaction is $SO_2(g) + 2H_2O(l) \rightarrow HSO_4^-(aq) + 3H^+(aq) + 2e^-(aq)$

The reduction half reaction is $MnO_4^-(aq) \rightarrow Mn^(2+)(aq)$.

In the reduction half reaction, the oxidation number of Mn changes from +7 to +2. Hence, 5 electrons are added to LHS of the reaction.

$MnO_4^-(aq) + 5e^- \rightarrow Mn^{2+}(aq)$

Charge is balanced in the reduction half reaction by adding 8 hydrogen ions to LHS.

$MnO_4^-(aq) + 5e^- + 8H^+(aq) \rightarrow Mn^{2+}(aq)$

To balance O atoms, 4 water molecules are added on RHS.

$MnO_4^-(aq) + 5e^- + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

To equalize the number of electrons, the oxidation half reaction is multiplied by 5 and the reduction half reaction is multiplied by 2.

$5SO_2(g) + 10H_2O(l) \rightarrow 5HSO_4^-(aq) + 15H^+(aq) + 10e^-(aq)$

$2MnO_4^-(aq) + 10e^- + 16H^+(aq) \rightarrow 2Mn^{2+}(aq)$

Two half cell reactions are added to obtain the balanced equation.

$2MnO_4^-(aq) + 5SO_2(g) + 2H_2O(l) \rightarrow 2Mn^{2+}(aq) + 5HSO_4^-(aq)$

#423677

Topic: Balance redox reactions

Passage

Balance the following redox reactions by ion electron method.

$H_2O_2(aq) + Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + H_2O(l)$(in acidic solution)

Solution

The oxidation half reaction is $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$.

The reduction half reaction is $H_2O_2(aq) + 2H^+(aq) + 2e^- \rightarrow 2H_2O(l)$

In above half reactions, all the atoms are balanced.

The oxidation half reaction is multiplied by 2 so as to balance the number of electrons in oxidation half reaction with reduction half reaction.

$2Fe^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + 2e^-$.

The oxidation half reaction is then added to the reduction half reaction to obtain balanced chemcial equation.

$H_2O_2(aq) + 2Fe^{2+}(aq) + 2H^+ \rightarrow 2Fe^{3+}(aq) + 2H_2O(l)$

#423679

Topic: Balance redox reactions

Passage

Balance the following redox reactions by ion electron method.

$Cr_2O_7^{2-} + SO_2(g) \rightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$(in acidic solution)

Solution

The oxidation half reaction is $SO_2(g) + 2H_2O(l) \rightarrow SO_4^{2-} + 4H^+(aq) + 2e^-$.

The reduction half reaction is $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$

The oxidation half reaction is multiplied by 3 and added to the reduction half reaction to obtain the balanced redox reaction.

$Cr_2O_7^{2-} + 3SO_2(g) + 2H^+(aq) \rightarrow 2Cr^{3+}(aq) + 3SO_4^{2-}(aq) + H_2O(l)$

#423686

Topic: Balance redox reactions

Passage

Balance the following equations in basic medium by ion-electron method and oxidation number methods and identify the oxidising agent and the reducing agent.

$P_4(s) + OH^-(aq) \rightarrow PH_3(g) + HPO_2^-(aq)$

Solution

Oxidation number method:

The oxidation number of P decreases from 0 to -3 and increases from 0 to +2. Hence, P_4 is oxidizing as well as reducing agent.

During reduction, the total decrease in the oxidation number for 4 P atoms is 12.

During oxidation, total increase inn teh oxidation number for 4 P atoms is 4.

The increase in the oxidation number is balanced with decrease in the oxidation number by multiplying $H_2PO_2^-$ with 3.

$P_4(s) + OH^-(aq) \rightarrow PH_3(g) + 3H_2PO_2^-(aq)$

To balance O atoms, multiply OH^- ions by 6.

$P_4(s) + 6OH^-(aq) \rightarrow PH_3(g) + 3H_2PO_2^-(aq)$

To balance H atoms, 3 water molecules are added to L.H.S and 3 hydroxide ions on R.H.S.

$P_4(s) + 6OH^-(aq) + 3H_2O(l) \rightarrow PH_3(g) + 3H_2PO_2^-(aq) + 3OH^-(aq)$

Subtract 3 hydroxide ions from both sides.

$P_4(s) + 3OH^-(aq) + 3H_2O(l) \rightarrow PH_3(g) + 3H_2PO_2^-(aq)$

Ion electron method:

The oxidation half reaction is $P_4(s) \rightarrow H_2PO_2^-(aq)$.

The P atom is balanced.

$P_4(s) \rightarrow 4H_2PO_2^-(aq)$

The oxidation number is balanced by adding 4 electrons on RHS.

$P_4(s) \rightarrow 4H_2PO_2^-(aq) + 4e^-$

The charge is balanced by adding 8 hydroxide ions on LHS.

$P_4(s) + 8OH^-(aq) \rightarrow 4H_2PO_2^-(aq)$

The O and H atoms are balanced.

The reduction half reaction is $P_4(s) \rightarrow PH_3(g)$.

The oxidation number is balanced by adding 12 eelctrons on LHS.

$P_4(s) + 12e^- \rightarrow PH_3(g)$

The charge is balanced by adding 12 hydroxide ions on RHS.

$P_4(s) + 12e^- \rightarrow PH_3(g) + 12OH^-$

The oxidation half reaction is multiplied by 3 and the reduction half reaction is multiplied by 2.

The half reactions are then added to obtain balanced chemical equation.

$P_4(s) + 3OH^-(aq) + 3H_2O(l) \rightarrow PH_3(g) + 3H_2PO_2^-(aq)$

#423687

Topic: Balance redox reactions

Passage

Balance the following equations in basic medium by ion-electron method and oxidation number methods and identify the oxidising agent and the reducing agent.

$N_2H_4(l) + ClO_3^-(aq) \rightarrow NO(g) + Cl^-(g)$

Solution

The oxidation number of N increases from -2 to $+2$. The oxidation number of Cl decreases from $+5$ to -1. Hence, hydrazine is the reducing agent and chlorate ion is the oxidizing agent.

Ion-electron method:

The oxidation half reaction is

$N_2H_4(l) \rightarrow NO(g)$

Balance the N atoms.

$N_2H_4(l) \rightarrow 2NO(g)$

To balance oxidation number, add 8 electrons.

$N_2H_4(l) \rightarrow 2NO(g) + 8e^-$

Add 8 hydroxide ions are to balance the charge.

$N_2H_4(l) + 8OH^-(aq) \rightarrow 2NO(g) + 8e^-$

The reduction half reaction is

$ClO_3^-(aq) \rightarrow Cl^-(aq)$

Add 6 electrons to balance the oxidation number.

$ClO_3^-(aq) + 6e^- \rightarrow Cl^-(aq)$

Add 6 hydroxide ions to balance the charge.

$ClO_3^-(aq) + 6e^- \rightarrow Cl^-(aq) + 6OH^-(aq)$

Multiply the oxidation half reaction by 3 and multiply the reduction half reaction by 2. Add two half reactions obtain the balanced chemical equation.

$3N_2H_4(aq) + 4ClO_3^-(aq) \rightarrow 6NO(s) + 4Cl^-(aq) + 6H_2O(aq)$

Oxidation number method:

Total decrease in oxidation number of N is 8. Total increase in the oxidation number of Cl is 6.

Multiply N_2H_4 with 3 and multiply ClO_3^- with 4.

$3N_2H_4(l) + 4ClO_3^-(aq) \rightarrow NO(g) + Cl^-(aq)$

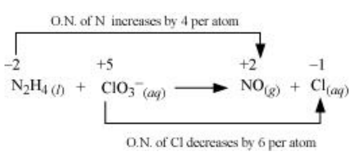
Balance N and Cl atoms.

$3N_2H_4(l) + 4ClO_3^-(aq) \rightarrow 6NO(g) + 4Cl^-(aq)$

Balance the O atoms are by adding 6 water molecules.

$3N_2H_4(l) + 4ClO_3^-(aq) \rightarrow 6NO(g) + 4Cl^-(aq) + 6H_2O(l)$

This is the balanced chemical equation.



#423688

Topic: Balance redox reactions

Passage

Balance the following equations in basic medium by ion-electron method and oxidation number methods and identify the oxidising agent and the reducing agent.

$Cl_2O_7(g) + H_2O_2(aq) \rightarrow ClO_2^-(aq) + O_2(g) + H^+$

Solution

The oxidation number of chlorine decreases from +7 to +3 and the oxidation number of O increases from -1 to zero. Thus, Cl_2O_7 is oxidizing agent and H_2O_2 is the reducing agent.

Ion electron method:

The oxidation half equation is

$H_2O_2(aq) \rightarrow O_2(g)$

To balance oxidation number, 2 electrons are added.

$H_2O_2(aq) \rightarrow O_2(g) + 2e^-$

2 hydroxide ions are added to balance the charge.

$H_2O_2(aq) + 2OH^- \rightarrow O_2(g) + 2e^-$

2 water molecules are added to balance the O atoms.

$H_2O_2(aq) + 2OH^- \rightarrow O_2(g) + 2H_2O(aq) + 2e^-$

The reduction half reaction is $Cl_2O_7 \rightarrow ClO_2^-(aq)$

The Cl atoms are balanced

$Cl_2O_7 \rightarrow ClO_2^-(aq)$

8 electrons are added to balance the oxidation number.

$Cl_2O_7 + 8e^- \rightarrow ClO_2^-(aq)$

6 hydroxide ions are added to balance the charge.

$Cl_2O_7 + 8e^- \rightarrow ClO_2^-(aq) + 6OH^-(aq)$

The oxidation half equation is multiplied with 4 and added to reduction half equation.

$CI_2O_7(g) + 4H_2O_2(aq) + 2OH^- \rightarrow CIO_2^-(aq) + 4O_2(g) + 5H_2O(l)$

Oxidation number method:

Total decrease in oxidation number of Cl_2O_7 is 8.

Total increase in oxidation number of H_2O_2 is 2.

H_2O_2 and O_2 are multiplied with 4

$Cl_2O_7(g) + 4H_2O_2(aq) \rightarrow ClO_2^-(aq) + 4O_2(g)$

Chlorine atoms are balanced

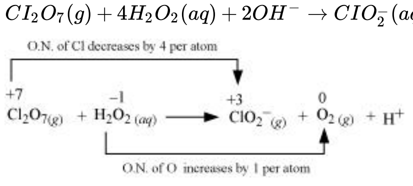
$Cl_2O_7(g) + 4H_2O_2(aq) \rightarrow 2ClO_2^-(aq) + 4O_2(g)$

O atoms are balanced by adding 3 water molecules.

$Cl_2O_7(g) + 4H_2O_2(aq) \rightarrow 2ClO_2^-(aq) + 4O_2(g) + 3H_2O(l)$

H atoms are balanced by adding 2 hydroxide ions and 2 water molecules.

$CI_2O_7(g) + 4H_2O_2(aq) + 2OH^- \rightarrow CIO_2^-(aq) + 4O_2(g) + 5H_2O(l)$



#423689

Topic: Types of redox reactions

What sorts of informations can you draw from the following reaction ?

$(CN)_2(g) + 2OH^-(aq) \rightarrow CN^-(aq) + CNO^-(aq) + H_2O(l)$

Solution

The reaction is a disproportionation reaction.

It occurs in basic medium.

The oxidation number of N in $(CN)_2$, CN^- and CNO^- is $-3, \ -2$and -5 respectively.

Cyanogen $(CN)_2$ is simultaneously reduced to CN^- ion and oxidised to cyanate ion CNO^- ion.

#423690

Topic: Balance redox reactions

The Mn^{3+} ion is unstable in solution and undergoes disproportionation to give $Mn^{2+}, \ MnO_2$ and H^+ ion. Write a balanced ionic equation for the reaction.

Solution

The unbalanced chemical reaction is:

$Mn^{3+}(aq) \rightarrow Mn^{2+}(aq) + MnO_2(s) + H^+(aq)$

The oxidation half reaction is,

$Mn^{3+}(aq) \rightarrow MnO_2(s)$.

To balance oxidation number, one electron is added on R.H.S.

$Mn^{3+}(aq) \rightarrow MnO_2(s) + e^-$

4 protons are added to balance the charge.

$Mn^{3+}(aq) \rightarrow MnO_2(s) + 4H^+(aq) + e^-$

2 water molecules are added to balance O atoms.

The reduction half reaction is $Mn^{3+}(aq) \rightarrow Mn^{2+}(aq)$.

An electron is added to balance oxidation number.

$Mn^{3+}(aq) + e^- \rightarrow Mn^{2+}(aq)$

Two half cell reactions are added to obtain balanced chemical equation.

$2Mn^{3+}(aq) + 2H_2O(l) \rightarrow Mn^{2+}(aq) + MnO_2(s) + 4H^+(aq)$

#423696

Topic: Balance redox reactions

Chlorine is used to purify drinking water. Excess of chlorine is harmful. The excess of chlorine is removed by treating with sulphur dioxide. Present a balanced equation for the redox change taking place in water.

Solution

The balanced chemical reaction for the redox reaction between chlorine and sulphur dioxide is:

$Cl_2 + SO_2 + 2H_2O \rightarrow 2Cl^- + SO_4^{2-} + 4H^+$.

#423698

Topic: Types of redox reactions

Passage

Refer to the periodic table given in your book and now answer the following questions:

Select three possible non metals that can show disproportionation reaction.

Solution

Phosphorous, chlorine and sulphur are the non metals which can show disproportionation reaction. The reactions shown by them are:

$P_4 + 3OH^- + 3H_2O \rightarrow PH_3 + 3H_2PO_2^-$

$Cl_2 + 2OH^- \rightarrow Cl^- + ClO^- + H_2O$

$S_8 + 12OH^- \rightarrow 4S^{2-} + 2S_2O_3^{2-} + 6H_2O$

#423699

Topic: Types of redox reactions

Passage

Refer to the periodic table given in your book and now answer the following questions:

Select three metals that can show disproportionation reaction.

Solution

Copper, gallium and indium are the metals that show disproportionation reaction. The reactions are shown below.

$2Cu^{+} \rightarrow Cu^{2+} + Cu$

$3Ga^{+} \rightarrow Ga^{3+} + 2Ga$

$3In+ \rightarrow In^{3+} + 2In$

#423724

Topic: Electrode potential

Passage

Depict the galvanic cell in which the reaction $Zn(s) + 2Ag^{+}(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$takes place. Further show:

the carriers of the current in the cell.

Solution

For the given redox reaction, the galvanic cell is

$Zn|Zn^{2+}(aq)||Ag^{+}(aq)|Ag$

The current is carried by the ions in the cell.